CHAPTER

ELECTROCHEMISTRY

MULTIPLE CHOICE QUESTIONS

1. **Symbol of hydronium ion is:**
   (a) H^+ (b) OH^- (c) H_3O^+ (d) none of these
2. **An example of non electrolyte is:**
   (a) glucose (b) aq NaCl (c) HCl (d) H_2SO_4
3. **An example of weak electrolyte is :**
   (a) HNO_3 (b) HCl (c) H_2SO_4 (d) H_2CO_3
4. **During electrolysis _____ takes place at anode.**
   (a) catenation (b) oxidation (c) reduction (d) addition
5. **Which is not the characteristic of electrolyte?**
   (a) cheap (b) conductor (c) easily oxidized (d) soluble in water
6. **In dry cell ___acts as cathode:**
   (a) Zn cup (b) graphite rod (c) paste (d) steel rod
7. **In Zn-Cu galvanic cell, Zn is dipped in:**
   (a) $ZnSO_4$ (b) $Zn(NO_3)_2$ (c) $CuSO_4$ (d) both a & b
8. **In Zn-Cu galvanic cell, Zn is used as:**
   (a) cathode (b) anode (c) electrode (d) all of above
9. **In $Zn + Cu^{+2} \longrightarrow Zn^{+2} + Cu$, Zn is:**
   (a) oxidized (b) reduced (c) redoxed (d) decomposed
10. **Anions are ________ions.**
    (a) positive (b) neutral (c) negative (d) amphoeteric
11. **Cations are _______ ions.**
    (a) negative (b) positive (c) neutral (d) none of these
12. **Electrolysis of NaCl is done in the cell:**
    (a) electrolytic (b) voltaic (c) down's (d) faradays'
13. **Which one is not produced during electrolysis of aqueous NaCl?**
    (a) Na^+ (b) OH^- (c) H_2O (d) Cl^-
14. **In pure water, out of ____only one molecule ionizes.**
    (a) 6×10^6 (b) 6×10^8 (c) 6×10^{16} (d) 6×10^{12}
15. **Which one is weak electrolyte?**
    (a) citric acid (b) carbonic acid (c) tartaric acid (d) all of these

16. **Metallic, ammonium and hydrogen ions carry _____ charges.**
(a) negative (b) neutral (c) positive (d) partial positive

17. **Electrode through which electrons enter the external circuit:**
(a) anode (b) Cathode (c) electrode (d) none of these

18. **Electrode through which electrons leave the external circuit:**
(a) anode (b) Cathode (c) graphite (d) electrolyte

19. **Which one is conductor?**
(a) naphthalene (b) paraffin wax (c) plastic (d) HCl

20. **Which solution is not a conductor:**
(a) NaCl (b) KI (c) $Co(NH_2)^2$ (d) $CuCl_2$

21. **Rods through which electric current enters or leaves the cell:**
(a) protons (b) electrons (c) electrodes (d) all of these

22. **Spontaneous redox reaction produce current in:**
(a) voltaic cell (b) electrolytic cell (c) galvanic cell (d) none of these

23. **In an oxidation reaction electrons are:**
(a) absorbed (b) lost (c) moved (d) increased

24. **In reduction reaction electrons are:**
(a) lost (b) absorbed (c) kept constant (d) all of these

ANSWER KEY

1	c	6	a	11	b	16	c	21	c
2	a	7	a	12	c	17	a	22	d
3	d	8	b	13	c	18	b	23	b
4	b	9	a	14	b	19	d	24	b
5	a	10	c	15	d	20	c	KIPS	

SHORT QUESTIONS

7.1 OXIDATION AND REDUCTION REACTIONS

Q.1 How can you justify that a reaction between magnesium and oxygen is a redox reaction, while the reaction shows only addition of oxygen (oxidation)

$2\,Mg + O_2 \longrightarrow 2\,MgO$

Ans.

(i) $2Mg \xrightarrow{Oxidation} 2Mg^{+2} + 4e^-$

Magnesium atom Magnesium ion

(ii) $O_2 + 4e^- \xrightarrow{Reduction} 2O^{-2}$

Oxygen molecule Oxide ion

(iii)

In above reaction, loss of electron also occur instead of addition of oxygen, while gain of electron is called reduction. So in redox reactions oxidation and reduction occur simultaneously.

Q.2 A reaction between carbon and oxygen involved only addition of oxygen (oxidation), but, it is called a redox reaction. Comment on this.

Ans. As the redox reaction

"A chemical reaction in which oxidation and reduction processes involved is called redox reaction."

In case of carbon reactys with oxygen:

$C \xrightarrow{Oxidation} C^{+4} + 4e^-$

Carbon atom Carbon ion

Oxidation reaction;

$O_2 + 4e^- \xrightarrow{Reduction} 2O^{-2}$

Oxygen molecule Oxide ions

$C^{+4} + 2O^{-2} \longrightarrow CO_2$

Loss of electrons is also called oxidation, while gain of electrons called reduction, so the above reaction is the example of redox reaction.

Q.3 Oxidation and reduction proceed simultaneously. Explain, with an example.

Ans. Oxidation and reduction occurs simultaneously in this reaction.

Example: $FeO + CO \longrightarrow Fe + CO_2$

Q.4 Identify which of the following is oxidation or reduction reaction.

(a) $K \longrightarrow K^+ + 1e^-$

(b) $Br + 1e^- \longrightarrow Br^-$

(c) $Cu \longrightarrow Cu^{+2} + 2e^-$

(d) $I^- \longrightarrow I + 1e^-$

(e) $Fe^{+2} \longrightarrow Fe^{+3} + 1e^-$

Ans. (a) Oxidation (Loss of electron)

(b) Reduction (Gain of electron)

(c) Oxidation (Loss of electron)

(d) Oxidation (Loss of electron)

(e) Oxidation (Loss of electron)

Q.5 An element M reacts with another element X to form MX_2. In terms of loss or gain of electrons, identify the element which is oxidized and which is reduced.

Ans. **Reaction:**

$$M + X_2 \longrightarrow MX_2$$

In this reaction, "M" release two electrons and become "M^{+2}" so oxidation reaction is:

$$M \xrightarrow{Oxidation} M^{+2} + 2e^-$$

While "X" gain two electrons, so reduction reaction is

$$2X + 2e^- \xrightarrow{Reduction} 2X^-$$

Overall reaction:

"M" is oxidized.

X is reduced.

Q.6 How can you justify that the following reaction is not only an oxidation reaction but also a complete redox reaction.

$$\mathbf{FeO + CO \longrightarrow Fe + CO_2}$$

Ans.

Ferrous oxide (FeO) transfer its oxygen to carbon monoxide (CO) to form carbon dioxide (CO_2) and Iron (Fe). In this reaction, oxidation and reduction take place simultaneously as:

$$\mathbf{FeO + CO \longrightarrow Fe + CO_2}$$

Q.7 Explain the term oxidation on the basis of electronic concept with an example

Ans. According to Electronic Concept:

"The loss of electrons by an atom or an ion is called oxidation."

Example:

$$Mg \xrightarrow{Oxidation} Mg^{+2} + 2e^-$$

7.2 OXIDATION STATE AND RULES FOR ASSIGNING OXIDATION STATE

Q.1 Find out the oxidation numbers of the following elements marked in bold in the formulae: $Ba_3(PO_4)_2$, $CaSO_4$, $Cu(NO_3)_2$, $Al_2(SO_4)_3$

Ans. **$Ba_3(PO_4)_2$:**

Barium sulphate

(Oxidation no. of "Ba") 3 + (Oxidation no. of "P")2 + (Oxidation no. of oxygen) 8 = 0

(+2)3 + (Oxidation no. of "P")2 + (-2)8 = 0

+6 + (Oxidation no. of "P")2 - 16 = 0

+6 + (Oxidation no. "P")2 = + 16

(Oxidation no. of "P" = +16 – 6 = 10

Oxidation no. of "P" $= \frac{+10}{2} = +5$

Q.4 In H_2S, SO_2 and H_2SO_4 the sulphur atom has different oxidation number. Find out the oxidation number of sulphur in each compound.

Ans.

(i) H_2S:

[O.N of H) x 2 + (O.N of "S") = 0
(+1) x 2 + (O.N of "S") = 0
2 + (O.N of "S") = 0
O.N of sulphur = -2

(ii) SO_2:

(O.N of sulphur) + (O.N of oxygen) × 2 = 0
(O.N of sulphur) + (-2) × 2 = 0
O.N of sulphur - 4 = 0
O.N of sulphur = +4

(iii) H_2SO_4:

(O.N of hydrogen) × 2 + (O.N of sulphur) + (O.N of oxygen) × 4 = 0
(O.N of "H") × 2 + (O.N. of "S") + (O.N of "O") × 4 = 0
(+1) 2 + (O.N of S) + (- 2) 4 = 0
+2 + (O.N of S) + (-2) 4 = 0
O.N of sulphur = + 8 -2
O.N of sulphur = +6

Q.5 An element X in oxidation state +7 gains electrons to be reduced to oxidation state +2. How many electrons did it accept?

Ans. Consider that element is "X"
As from equation:
$X^{+7} + 5e^- \longrightarrow X^{+2}$
This element will gain 5 electron to be reduced.

Q.6 An element "X" has oxidation state "0". What will be its oxidation state when it gains three electrons?

Ans. $X_o + 3e^- \longrightarrow X^{-3}$
Its oxidation state will be (-3)

Q.7 If the oxidation state of an element changes from +5 to – 3, Has it been reduced or oxidized? How many electrons are involved in this process?

Ans. For example the element is "X". As the reaction:
$X^{+5} + 8e^- \longrightarrow X^{-3}$
This element gain the electrons, so reduction will be occurred.
It will gain 8 electron 5 are used to neutral its oxidation state while due to more 3 electron it will attain -3 charge.

7.3 OXIDIZING AND REDUCING AGENTS

Q.1 In the following reaction, how can you justify that H_2S is oxidized and SO_2 is reduced.

$SO_2 + 2H_2S \longrightarrow 2H_2O + 3S$

Ans. in this reaction, H_2S release the electrons while SO2 absorb the electrons as:

Q.2 The reaction between Mn02 and HCl is a redox reaction written as balance chemical equation.

$$MnO_2 + 4HCl \longrightarrow MnCl_2 + 2H_2O + Cl_2$$

Find out:

(a) The substance oxidized

Ans. MnO_2 is oxidized and it lose electron.

(b) The substance reduce.

Ans. Chlorine is reduced, it gain electrons.

(c) The substance reduced

Ans. Cl_2 is acts oxidizing agent.

(d) The substance which acts as reducing agent.

Ans. MnO2 acts as reducing agent.

Q.3 The following reactions are redox reaction. Find out the element which has been reduced and the element which has been oxidized?

(a) $Zn + CuSO_4 \longrightarrow ZnSO_4 + Cu$

(b) $Cu + 2AgNO_3 \longrightarrow Cu(NO_3)_2 + 2Ag$

(c) $H_2S + Cl_2 \longrightarrow 2HCl + S$

Ans. **(a)** Zn is oxidized and Cu is reduced.

(b) Cu is oxidized and Ag is reduced.

(c) S is oxidized and Cl is reduced.

Q.4 Why the following reaction is not a redox reaction. Explain with reasons?

$NaOH + HCl \longrightarrow NaCl + H_2O$

Ans. In this reaction, as

$$NaOH + HCl \longrightarrow NaCl + H_2O$$

There is no oxidation reduction processes takes place. It is an acid base reaction. Non of atom oxidized or reduced, in this reaction.

7.5 ELECTRO CHEMICAL CELLS

Q.1 Why are the strong electrolytes termed as good conductors?

Ans. Strong electrolytes are good conductor because it completely ionized into its ions and it readily produced more ions as:

$HCl_{(aq)} \longrightarrow H^+_{(aq)} + Cl^-_{(aq)}$ and $NaOH_{(aq)} \longrightarrow Na^+_{(aq)} + OH^-_{(aq)}$

Q.2 Does non-electrolytes forms ions in solution?

Ans. Non-electrolytes does not form ions in solution.

Q.3 What is difference between a strong electrolyte and a weak electrolyte?

Ans.

Strong Electrolyte	Weak Electrolyte
Definition: The electrolytes which ionize completely in solution and produce more ions, are called strong electrolytes.	**Definition:** The electrolytes which ionize to a small extent when dissolve in water and could not produce more ions are called weak electrolytes.
Example: Strong electrolytes are aqueous solutions of NaCl, NaOH and H_2SO_4. $NaOH_{(s)} \xrightarrow{H_2O} Na^+_{(aq)} + OH^-_{(aq)}$	**Example:** Weak electrolytes are the aqueous solution of acetic and or $Ca(OH)_2$ $CH_3COOH_{(l)} + H_2O_{(l)} \longrightarrow CH_3COO^-_{(aq)} + H_3O^+_{(aq)}$

Q.4

Q.5 Identify a strong or weak electrolyte among the following compounds:

$CUSO_4$, H_2CO_3, $Ca(OH)_2$, HCI, $AgNO_3$

Ans. $AgNO_3$ and $CuSO_2$ are strong electrolytes while H_2CO_3 and $Ca(OH)_2$ are the example of weak electrolytes.

Q.6 Which force drives the non-spontaneous reaction to take place?

Ans. Non-spontaneous reactions take place in the presence of an external agent. That external agent are electrons that cause electricity. So electric energy helps the non-spontaneous reactions to proceed.

Q.7 Which type of chemical reaction takes place in electrolytic cell?

Ans. In electrolytic cell, non-spontaneous reaction takes place.

Q.8 What type of reaction takes place at anode in electrolytic cell?

Ans. Oxidation takes place at anode. Anode is a positive charge electrode. The atoms on this electrode release electrons as:

$$2Cl^- \longrightarrow 2Cl + 2e^-$$

Q.9 Why the positively charged electrode is called anode in electrolytic cell?

Ans. The positive charged electrode is called anode because it connected to the (+) terminal of the battery and all electron move towards it as

$$2Cl^- \longrightarrow 2Cl + 2e$$

Q.10 In the electrolysis of water, towards which terminal hydronium ions move?

Ans. In the electrolysis of water, H_3O (hydronium) + or H^+ ions move toward the cathode.

Reaction at cathode: $H^+ + 1e^- \longrightarrow [H]$

Q.11 In the electrolysis of water, where is the oxygen produced?

Ans. In the electrolysis of water oxygen is produced at anode as

Reaction at anode: $4OH^- \longrightarrow 2H_2O_{(aq)} + O_{2(g)} + 4e-$

Q.12 Towards which electrode of the electrolytic cell moves the cations and what does they do there?

Ans. Cation ions carry (+) charge, they move towards the cathode, in an electrolytic cell. They gain electron at cathode and oxidized.

Q.13 How the half cells of a galvanic cell are connected? What is function of salt bridge?

Ans. The two half cells of the galvanic cell are connected by salt bridge. Salt bridge is used to maintain the flow of ionic current and the electric neutrality.

7.6 ELECTRO CHEMICAL INDUSTRIES

Q.1 Anode of Downs cell is made of a non-metal, what is its name? What is the function of this anode?

Ans. In Downs cell, there is a large block of graphite, which acts as an anode. The Cl^- ions are oxidized there produce Cl_2 gas at anode.

Q.2 Where does the sodium metal is collected in Downs cell?

Ans. Na+ are reduced at cathode and molten Na-metal floats on the denser molten salt mixture from where it is collected in a side tube.

Q.3 What is the name of the by-product produced in the Downs cell?

Ans. Hydrogen gas and chlorine gas are the by-products produced in the Downs cell.

Q.4 Are anodes of Downs cell and Nelson cell made of same element? If yes, what is its name?

Ans. Yes, in both cells anode is made of graphite.

Q.5 What is the shape of cathode in Nelson's cell? Why is it perforated?

Ans. In Nelson's cell, cathode is made of iron and have U-shaped. It is perforated because extra brine or NaOH can be easily separated out.

Q.6 Which 'ions are discharged at cathode in Nelson's cell and what is produced at cathode?

Ans. The H+ ions are discharged at cathode in Nelson's cell. At cathode, sodium hydroxide (NaOH) and hydrogen (H_2) are produced.

7.7 CORROSION AND ITS PREVENTION

Q.1 What is the difference between corrosion and rusting?

Ans.

Corrosion	Rusting
• Corrosion is the general term used for all the metals	• Corrosion of iron is called rusting
• Corrosion of some metals may be stopped	• Rusting is the continuous process.
• It is redox reaction	• It is also redox reaction.

Q.2 What happens to iron in the rusting process?

Ans. During rusting Fe^{+2} formed that spreads through out the surrounding water and react with O_2 to form the salt $Fe_2O_3.nH_2O$ which is called rust.

Q.3 Rusting completes in how many redox reactions?

Ans. Rusting completes in two redox reactions as:

First Redox Reaction:

$$2Fe_{(s)} + O_{2(aq)} + 4H^{+}_{(aq)} \longrightarrow 2Fe^{+2} + 2H_2O$$

Second Redox Reaction:

$$2Fe^{+2} + \frac{1}{2}O_2 + (2+n)H_2O \longrightarrow Fe_2O_3.nH_2O + 4H^{+}$$

Q.4 Explain the role of O_2 in rusting?

Ans. Oxygen plays important role in rusting. Electrons reduce the oxygen molecules in the presence of H+ ions:

Q.5 State the best method for protection of metal from corrosion.

Ans. The best method for protection of metal from corrosion is the coating of highly resistant metal. Corrosion resistant metals like Zn, Sn and Cr are coated on the surface of iron to protect it from corrosion.

Q.6 What do you mean by galvanizing?

Ans. The process of coating a thin layer of zinc on iron is called galvanizing.

Q.7 What is the advantage of galvanizing?

Ans. A big advantage of galvanizing is that zinc protects the iron against corrosion even after the coating surface is broken.

Q.8 Why tin plated iron is rusted rapidly when tin layer is broken?

Ans. When tin layer is broken the iron is exposed to the air and water, a galvanic cell is established and iron rusts rapidly.

Q.9 Name the metal which is used for galvanizing iron?

Ans. Zinc metal is used for galvanizing iron.

7.8 ELECTROPLATING

Q.1 Define electroplating?

Ans. Electroplating is depositing of one metal over the other by means of electrolysis.

Q.2 How electroplating of zinc is carried out?

Ans. The zinc is deposited on the metal by immersing it in a chemical bath containing electrolyte zinc sulphate. A current is applied which results in zinc being deposited on the target metal i.e., cathode.

Q.3 Which material is used to make cathode electroplating?

Ans. The cathode is made of the objects that is to be electroplated e.g., sheet of iron.

Q.4 Why is the anode made up of a metal to be deposited during electrolysis

Ans. Because, when the current passed, the metal from anode dissolved in the solution and metallic ions migrate to the cathode and discharge or deposit on the object. As a result of this charge, a thin layer of metal deposits on the object.

LONG QUESTIONS

INTRODUCTION

Q.No.1 What is Electrochemistry? Write down the types of redox reaction.

Definition

Electrochemistry is the branch of Chemistry that deals with the relations between electricity and chemical reactions.

Explanation

Electrochemistry involves oxidation and reduction reactions, which are also known as redox reactions.

Types of Redox reactions

There are two types of redox reaction

(i) Spontaneous reaction

(ii) Non-spontaneous reactions

Spontaneous reaction

Spontaneous reactions are those which take place on their own without an external agent

Non-spontaneous reactions

Non-spontaneous reactions are those which take place in the presence of an external agent. These reactions take place in galvanic or electrolytic cells.

Q.No.2 Explain the concept of Oxidation and Reduction Reactions.

One concept of oxidation and reduction is based upon either addition or removal of oxygen or addition or removal of hydrogen or addition or removal of electrons in a chemical reaction.

OXIDATION AND REDUCTION REACTIONS IN TERM OF LOSS OR GAIN OF HYDROGEN/OXYGEN

Oxidation

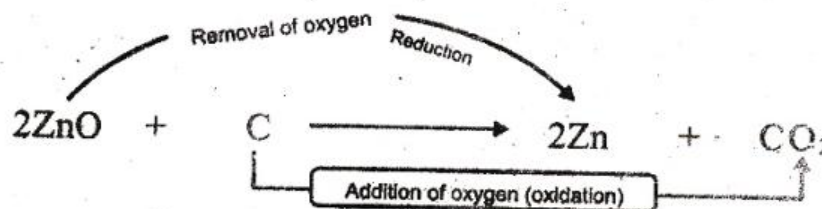
Oxidation is defined as addition of oxygen or removal of hydrogen during a chemical reaction.

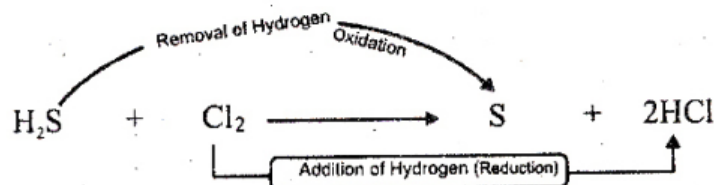
Reduction

Reduction is defined as addition of hydrogen or removal of oxygen during a chemical reaction. Both of these processes take place simultaneously in a reaction.

Examples

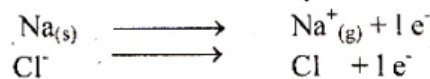
(i) A reaction between zinc oxide and carbon takes place by the removal of oxygen (reduction) from zinc oxide and addition of oxygen (oxidation) to carbon.

$$ZnO + C \longrightarrow Zn + CO$$

(ii) A reaction between hydrogen sulphide and chlorine takes place by oxidation of hydrogen sulphide and reduction of chlorine. Hydrogen is being removed from H_2S and added to chlorine.

$$H_2S + Cl_2 \longrightarrow HCl + S$$

OXIDATION AND REDUCTION IN TERMS OF LOSS OR GAIN OF ELECTRON

Oxidation

Oxidation is loss of electrons by an atom or an ion.

$$Na_{(s)} \longrightarrow Na^+_{(g)} + 1e^-$$
$$Cl^- \longrightarrow Cl + 1e^-$$

Reduction

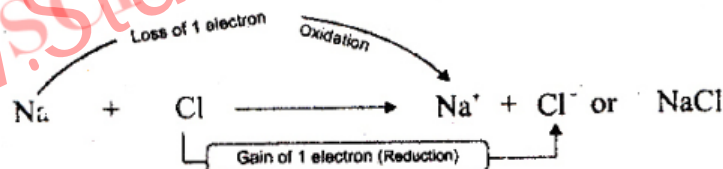
Reduction is gain of electrons by an atom or ion. e.g.

$$2H^+ + 2e^- \longrightarrow H_2$$
$$Cl_2 + 2e^- \longrightarrow 2Cl^-$$

Explanation

A reaction between sodium metal and chlorine takes place in three steps.

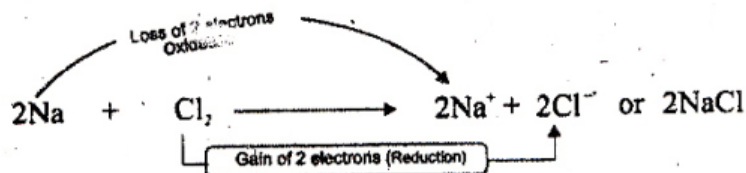
(i) First sodium atom losses an electron, to form sodium ion.

$$Na \longrightarrow Na + 1e^- \quad \text{(oxidation)}$$

(ii) Simultaneously, this electron is accepted by chlorine atom (reduction process), as chlorine atom needs one electron to complete its octet. As a result chlorine atom changes to chloride ion.

$$Cl\,(g) + 1e^- \longrightarrow Cl^- \text{ (reduction)}$$

Loss of 2 electrons Oxidation

$$2Na + Cl_2 \longrightarrow 2Na^+ + 2Cl^- \text{ or } 2NaCl$$

Gain of 2 electrons (Reduction)

(iii) Ultimately, both these ions attract each other to form sodium chloride. Complete redox reaction is sum of the oxidation and reduction reactions between sodium and chlorine atoms and it is represented as;

$$Na^+ + Cl^- \longrightarrow NaCl$$

Summary of concept of oxidation and reduction

Oxidation	Reduction
Addition of oxygen	Removal of oxygen
Removal of hydrogen	Addition of hydrogen
Loss of electrons	Gain of elections

Q.No.3 Define oxidation state and explain the rules used for assigning oxidation state.

Oxidation state

Oxidation state or oxidation number (O.N.) is the apparent charge assigned to an atom of an element in a molecule or in an ion.

Example

In HCl, oxidation number of H is + 1 and that of Cl is -1.

Rules for assigning oxidation state

(i) The oxidation number of all elements in the free state is zero.

(ii) The oxidation number of an ion consisting of a single element is the same as the charge on the ion.

(iii) The oxidation number of different elements in the periodic table is: in Group-I it is +1, in Group-2 it is +2 and in Group-3 it is +3.

(iv) The oxidation number of hydrogen in all its compounds is + 1. But in metal hydrides it is -1.

(v) The oxidation number of oxygen in all its compounds is -2. But it is -1 in peroxides and +2 in OF_2.

(vi) In any substance the more electronegative atom has the negative oxidation number.

(vii) In neutral molecules, the algebraic sum of the oxidation numbers of all the elements is zero.

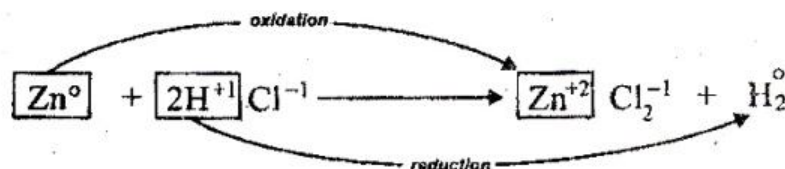
(viii) In ions, the algebraic sum of oxidation number equals the charge on the ion.

Q.No.4 Explain oxidizing and reducing agents with example.

Oxidizing agent

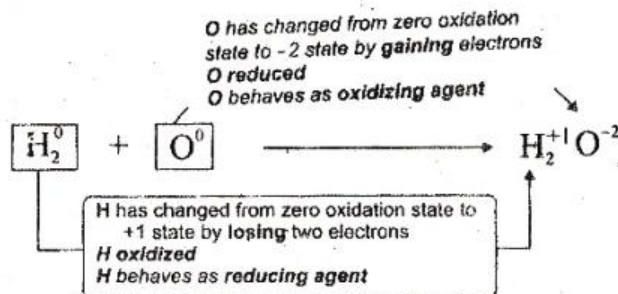
An oxidizing agent is the species that oxidizes a substance by taking electrons from it. The substance (atom or ion) which is reduced itself by gaining electrons is also called oxidizing agent.

Example

Non-metals are oxidizing agents because they accept electrons being more electronegative elements.

Reducing agent

Reducing agent is the species that reduces a substance by donating electron . to it. The substance (atom or ion) which is oxidized by losing electrons is also called reducing agent.

Example

Almost all metals are good reducing agents because they have the tendency to lose electrons.

Q.No.5 What are electrochemical cells? Explain the concept of electrolytes.

Electrochemical cell

Electrochemical cell is a system in which two electrodes are dipped in the solution of an electrolyte which are connected to the battery.

Electrochemical cell is an energy storage device in which either a chemical reaction takes place by using electric current (electrolysis) or chemical reaction produces electric current.

Types of electrochemical cells

(i) Electrolytic cells **(ii)** Galvanic cells

CONCEPT OF ELECTROLYTES

Electrolytes

The substances, which can conduct electricity in their solutions or molten states, are called electrolytes. For example, solutions of salts, acids or bases are good electrolytes. The electricity cannot pass through solid NaCl but its aqueous solution or molten NaCl are good electrolytes.

Classification of electrolytes

Electrolytes are classified into two groups depending upon their extent of ionization in solution.

Strong Electrolytes

The electrolytes which ionize completely in solution and produce more ions, are called strong electrolytes.

Example

Strong electrolytes are aqueous solutions of NaCl, NaOH and H_2SO_4 etc.

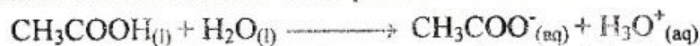
$$NaOH_{(s)} \longrightarrow Na^+ + {}^-OH$$

Weak electrolytes

The substances which ionize to a small extent when dissolved in water and could not produce more ions are called weak electrolytes.

Example

Acetic acid (CH_3COOH), and $Ca(OH)_2$ when dissolved in water, ionizes to a small extent and are good examples of weak electrolytes. Weak electrolytes do not ionize completely. For example, ionization of acetic acid in water produces less ions:

$$CH_3COOH_{(l)} + H_2O_{(l)} \longrightarrow CH_3COO^-_{(aq)} + H_3O^+_{(aq)}$$

As a result the weak electrolyte is a poor conductor of electricity.

Non-Electrolytes

The substances, which do not ionize in solution and do not allow the current to pass through their solutions, are called non-electrolytes.

Example

Sugar solution and benzene are non-electrolytes.

Q.No.6 What is electrolysis? Writes a note on Electrolytic cells.

ELECTROLYSIS

Definition

The chemical decomposition of a compound into its components by passing current through the solution of the compound or in the molten state of the compound is called electrolysis.

ELECTROLYTIC CELL

Definition

The type of electrochemical cell in which a non-spontaneous chemical reaction takes place when electric current is passed through the solution, is called an electrolytic cell.

Example

Downs cell, Nelson's cell

Construction of an Electrolytic Cell

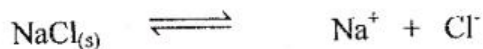
An electrolytic cell consists of a solution of an electrolyte, two electrodes (anode and cathode) that are dipped in the electrolytic solution and connected to the battery. The electrode connected to positive terminal is called anode and electrode connected to the negative terminal is called cathode.

Working of an Electrolytic Cell

When electric current is applied from battery, the ions in the electrolyte migrate to their respective electrodes. The anions, which are negatively charged, move towards the anode and discharge there by losing their electrons. Thus oxidation takes place at anode. While cations, which are positively charged ions, move towards cathode. Cations gain electrons from the cathode and as a result reduction takes place at cathode.

Example

When fused salt of sodium chloride is electrolyzed the following reactions take place during this process:

Ionization reaction

$$NaCl_{(s)} \rightleftharpoons Na^{+} + Cl^{-}$$

Oxidation reaction at anode

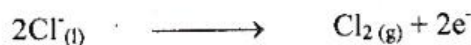
$$2Cl^{-}_{(l)} \longrightarrow Cl_{2\,(g)} + 2e^{-}$$

Reduction reaction at cathode

$$2\,Na^{+}_{(l)} + 2e^{-} \longrightarrow 2\,Na$$

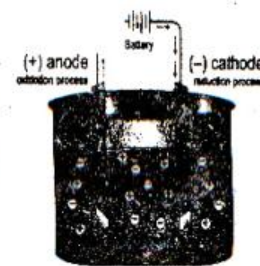
Overall reaction

$$2Na^{+} + 2e^{-} \longrightarrow 2Na$$
$$2Cl^{-}_{(l)} \longrightarrow Cl_{2\,(g)} + 2e^{-}$$

$$Na^{+} + Cl^{-} \longrightarrow 2Na + Cl_{2}$$

Q.No.7 Write a note on the electrolysis of water

Introduction

Pure water is a very weak electrolyte. It ionizes to a very small extent. The concentrations of hydrogen ions (H^+) and hydroxide ions (^-OH) are both at 10^{-7} mol dm^{-3} respectively. When a few drops of an acid are added in water, its conductivity improves.

$$H_2O \longrightarrow H^+ + {}^-OH$$

Electrolysis

When an electric current is passed through this acidified water, OH^- **anions** move towards positive electrode (anode) and H+ **cations** move towards negative electrode (cathode) and discharge takes place at these electrodes. They produce oxygen and hydrogen gases respectively at anode and cathode.

Oxidation reaction at anode

$$4OH^- \longrightarrow 2H_2O + O_2 + 4e-$$

Reduction reaction at cathode

$$4H^+ + 4e \longrightarrow 4H$$

$$4H \longrightarrow 2H_2$$

Overall reaction

$$4\,OH^- \longrightarrow 2H_2O + O_2 + 2H_2$$

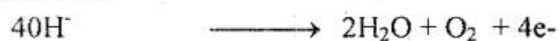
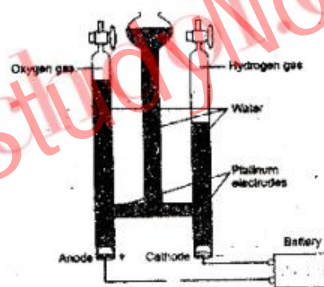



Fig. 7.2 Elect. .lytic cell showing electrolysis of water

Q.No.8 Write a detail note o. Galvanic Cell.

Definition

The electrochemical cell in which a spontaneous chemical reaction tak· · place and generates electric current is called Galvanic or voltaic cell.

EXAMPLE

Daniel cell

A volta (1745-1827) was an Italian physicist known especially for the development of the first electric cell in 1800.

Construction of a Daniel Cell

A galvanic cell consists of two cells, each called as half cell, connected electrically by a salt-bridge. In each of the half-cell, an electrode is dipped in 1M solution of its own salt and connected through a wire to an external circuit.

Left half cell

The left half cell consists of an electrode of zinc metal dipped in 1 M solution of zinc sulphate.

Right half

The right half cell is a copper electrode dipped in 1M solution of copper sulphate.

Salt bridge

Salt bridge is a U shaped glass tube. It consists of saturated solution of strong electrolyte supported in a jelly type material. The ends of the U tube are sealed with a porous material like glass wool.

Function of the salt bridge

To keep the solutions of two half cells neutral by providing a pathway for migration of Ions.

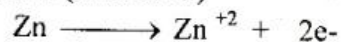
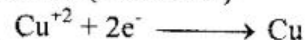
Working of the Cell

The Zn metal has tendency to lose electrons more readily than copper. As a result oxidation takes place at Zn-electrode. The electrons flow from Zn-electrode through the external wire in a circuit to copper electrode. These electrons are gained by the copper ions of the solution and copper atoms deposit at the electrode. The respective oxidation and reduction processes going on at two electrodes are as follows:

Reaction at anode (oxidation)

$$Zn \longrightarrow Zn^{+2} + 2e-$$

Reaction at cathode (reduction)

$$Cu^{+2} + 2e^{-} \longrightarrow Cu$$

Overall galvanic reaction

$$Zn_{(s)} + CU^{+2}_{(aq)} \longrightarrow Zn^{+2}_{(aq)} + Cu_{(s)}$$

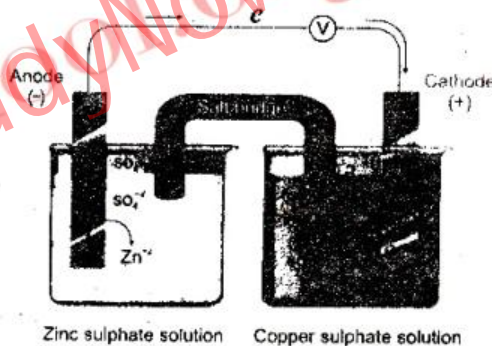
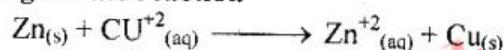



Fig. 7.3 A Daniel Cell

Uses

The batteries which are used for starting automobiles, running calculators and toys.

Q.No.9 Differentiate between electrolytic and galvanic cell

A Comparison of Electrolytic and Galvanic Cells

Electrolytic cell	Galvanic cell
It consists of one complete cell, connected to a battery.	It consist of two half cells connected. through a salt bridge.
Anode has positive charge while cathode was negative charge.	Anode has negative charge while cathode has positive charge.
Electrical energy is converted into chemical energy.	Chemical energy converted into electrical energy.
Currents is use for a non-spontaneous chemical reaction to take a place.	Redox reaction take place spontaneously and products electric current.
Example: Down cell	**Example:** Dray cell

Q.No.10 How sodium chloride is manufactured on industrial scale?

Manufacture of Sodium Metal from Fused NaCl

On the industrial scale molten sodium metal is obtained by the electrolysis of fused NaCl in the Downs cell. This electrolytic cell is a circular furnace. In the center there is a large block of graphite, which acts as an anode while cathode around it is made of iron.

Working of Downs Cell

The fused NaCl produces Na^+ and Cl^- ions, which migrate to their respective electrodes on the passage of electric current. The electrodes are separated by steel gauze to prevent the contact between the products. The Cl ions are oxidized to give Cl gas at the anode. It is collected over the anode within an inverted cone-shaped structure. While Na+ are reduced at cathode and molten Na metal floats on the denser molten salt mixture from where it is collected in a side tube. Following reactions take place during the electrolysis of the molten sodium chloride:

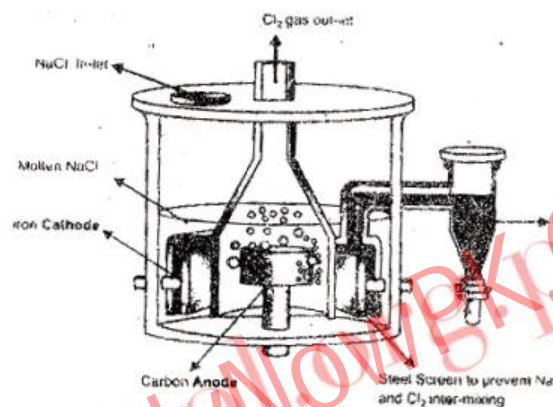



Fig. 7.4 Downs Cell for production of Sodium Metal

Ionization of molten NaCl

$$2NaCl \longrightarrow 2Na^+ + 2Cl^-$$

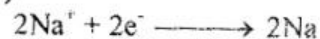
Reaction at anode (oxidation)

$$2Cl^- \longrightarrow Cl_2 + 2e^-$$

Reaction at cathode (reduction)

$$2Na^+ + 2e^- \longrightarrow 2Na$$

Overall reaction

$$2NaCl \longrightarrow 2Na + Cl_2$$

Q.No.11 Write a note on Nelson's Cell.

Introduction

On industrial scale caustic soda, sodium hydroxide NaOH, is produced in Nelson's cell by the electrolysis of aqueous solution of NaCl called brine.

Construction

It consists of a steel tank in which graphite anode is suspended in the center of a U shaped perforated iron cathode. This iron cathode is internally lined with asbestos diaphragm. Electrolyte brine is present inside the iron cathode.

Working of Nelson's Cell

Aqueous solution of sodium chloride consists of Na^+, Cl^-, H^+ and ^-OH ions. These ions move towards their respective electrodes and redox reactions take place at these electrodes. When electrolysis takes place Cl ions are discharged at anode and Cl_2 gas rises into the dome at the top of the cell. The H^+ ions are discharged at cathode and H_2 gas escapes through a pipe. The sodium hydroxide solution slowly percolates into a catch basin.

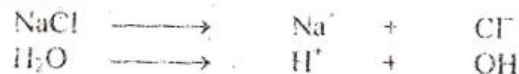
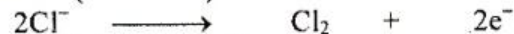
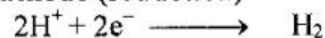
Ionization of Brine

$$NaCl \longrightarrow Na^+ + Cl^-$$
$$H_2O \longrightarrow H^+ + OH^-$$

Reaction at anode (oxidation)

$$2Cl^- \longrightarrow Cl_2 + 2e^-$$

Reaction at cathode (reduction)

$$2H^+ + 2e^- \longrightarrow H_2$$

Overall cell reaction of this process

$$2NaCl + 2H_2O \longrightarrow H_2 + Cl_2 + 2NaOH$$

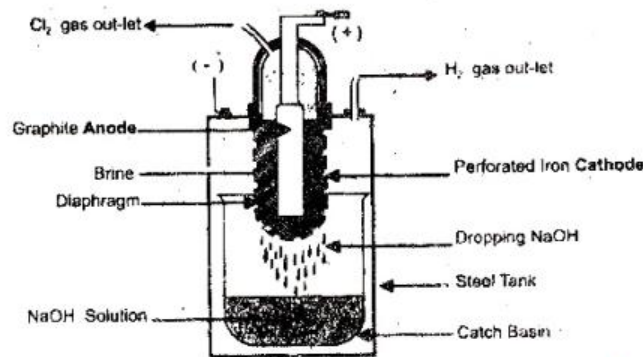
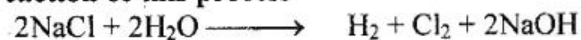



Fig. 7.5 Nelson's Cell for production of NaOH

Q.No.12 What is corrosion? How iron gets rust?

Corrosion

Corrosion is a redox chemical reaction that takes place by the action of air and moisture with the metals. The most common example of corrosion is rusting of iron.

Rusting of Iron

Corrosion is a general term but corrosion of iron is called rusting.

Conditions for rusting

The important condition for rusting is moist air (air having water vapours in it). There will be no rusting in water vapours free of air or air free of water.

Process of rusting

Stains and dents on the surface of the iron provide the sites for this process to occur. This region is called anodic region and following oxidation reaction takes place here:

$$2Fe \longrightarrow 2Fe^{+2} + 4e^-$$

This loss of electrons damages the object. The free electrons move through iron sheet ,until they reach to a region of relatively high O_2 concentration near the surface surrounded by water layer. This region acts as cathode and electrons reduce the oxygen molecule in the presence of H+ ions:

The overall redox process is completed without the formation of rust

$$2Fe + O_2 + 4H^+ \longrightarrow 2Fe^{+2} + 2H_2O$$

The Fe^{+2} formed spreads through out the surrounding water and react with O_2 to form the salt $Fe_2O_3 .nH_2O$ which is called **rust.** It is also a redox reaction.
The rust layer of iron is porous and does not prevent further corrosion. Thus rusting continues until all the piece of iron is eaten up

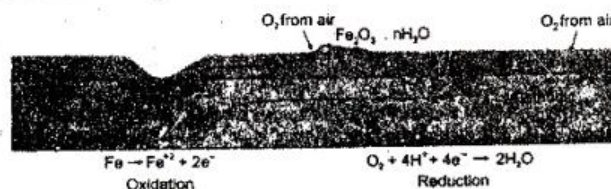



Fig. 7.6 Rusting of iron.

Q.No.13 How corrosion of metal can be avoided?

(i) Removal of stains

The regions of stains in an iron rod act as the site for corrosion. If the surface of iron is properly cleaned and stains are removed. It would prevent corrosion.

(ii) Paints and greasing

Polishing or painting of the surface can prevent the corrosion of iron. With development of technologies, modern paints contain a combination of chemicals called stabilizers that provide protection against the corrosion in addition to prevention against the weathering and other atmospheric effects.

(iii) Alloying

Alloy is a homogeneous mixture of one metal with one or more other metals or non-metals. Alloying of iron with other metals has proved to be very successful technique against rusting. The best example of alloying is the 'stainless steel', which is a good combination of iron, chromium and nickel.

(iv) Metallic coating

The best method for protection against the corrosion of metals exposed to acidic conditions is coating the metal. Corrosion resistant metals like Zn, Sn and Cr are coated on the surface of iron to protect it from corrosion. It is the most widely applied technique in the food industry where food is 'tin-packed'. The containers of iron are coated with tin or chromium to give it a longer life. Metallic coating can take place by physical as well as electrolytic methods.

(v) Physical Methods

(a) Zinc coating or Galvanizing

The process of coating a thin layer of zinc on iron is called galvanizing. This process is carried out by dipping a clean iron sheet in a zinc chloride bath and then heating it. After this iron sheet is removed, rolled into molten zinc metal bath and finally air-cooled. Advantage of galvanizing is that zinc protects the iron against corrosion even after the coating surface is broken.

(b) Tin Coating

It involves the dipping of the clean sheet of iron in a bath of molten tin and then passing it through hot pairs of rollers. Such sheets are used in the beverage and food cans. The tin protects the iron only as long as its protective layer remains intact. Once it is broken and the iron is exposed to the air and water, a galvanic cell is established and iron rusts rapidly.

Q.No.14 What is electroplating? Write down its procedure.

(vi) Electrolytic method

Definition

Electroplating is depositing of one metal over the other by means of electrolysis.

Purpose of electroplating

This process is used to protect metals against corrosion and to improve their appearance.

Principle of electroplating

Principle of electroplating is to establish an electrolytic cell in which anode is made of [illegible] to be deposited and cathode of the object on which metal is to deposit. The electrolyte [illegible]us solution of a salt of the respective metal.

Procedure for Electroplating

In this process the object to be electroplated is cleaned with sand and washed with caustic soda solution. The anode is made of the metal, which is to be deposited like Cr, Ni. The cathode is made up of the object that is to be electroplated like some sheet made up of iron. The electrolyte in this system is a salt of the metal being deposited. The electrolytic tank is made of cement, glass or wood in which anode and cathode are suspended. These electrodes are connected with a battery. When the current is passed, the metal from anode dissolves in the solution and metallic ions migrate to the cathode and discharge or deposit on the cathode (object). As a result of this discharge, a thin layer of metal deposits on the object, which is then pulled out and cleaned. Some examples of electroplating are discussed here.

Q.No.15 Write down the types of electroplating.

(a) Electroplating of Silver

The electroplating of silver is carried out by establishing an electrolytic cell. The pure piece of silver strip acts as anode that is dipped in silver nitrate solution. The cathode is the metallic object to be coated such as silver spoon. When the current is passed through the cell, the Ag^+ ions dissolve at the anode, and migrate towards the cathode where they discharge and deposit on the object e.g. spoon. The chemical reaction can be represented as:

At anode: $Ag_{(s)} \longrightarrow Ag^+_{(aq)} + 1e^-$

At cathode: $Ag^+_{(aq)} + 1e^- \longrightarrow Ag_{(s)}$

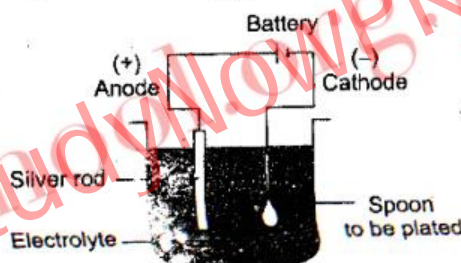
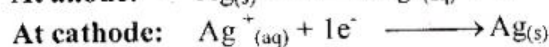
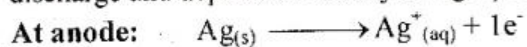
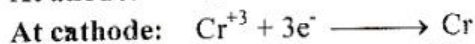
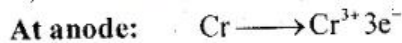



Fig. 7.7 Electroplating of an object.

(b) Electroplating of Chromium

The electroplating of chromium is carried out in the same way as that of silver. The object to be electroplated is dipped in aqueous solution of chromium sulphate containing a little sulphuric acid, that acts as an electrolyte. The object to be electroplated acts as cathode while anode is made of antimonial-lead. The electrolyte ionizes and provides Cr^{3+} ions, which reduce and deposit at cathode.

At anode: $Cr \longrightarrow Cr^{3+} 3e^-$

At cathode: $Cr^{+3} + 3e^- \longrightarrow Cr$

For practical convenience, the steel is usually plated first with nickel or copper and then by chromium because it does not adhere well on the steel surface. Moreover, it allows moisture to pass through it and metal is stripped off. The nickel or copper provides adhesion and then chromium deposited over the adhesive layer of copper lasts longer. This type of electroplating resists corrosion and gives a bright silvery appearance to the object.

(c) Deposition of zinc

The target metal is cleaned in alkaline detergent type solutions, and it is treated with acid, in order to remove any rust or surface scales. Next, the zinc is deposited on the metal by immersing it in a chemical bath containing electrolyte zinc sulphate. A current is applied, which results in zinc being deposited on the target metal i.e. cathode.

(d) Tin coating

Tin is usually ele troplated on steel by placing the steel into a container containing a solution of tin sa:. The teel is connected to an electrical circuit, acting as cathode. While the other electrode made of tin metal acts as anode. When an electrical current passes through the circuit, tin metal ions present in the solution deposit on steel.

(e) Electrolytic refining of copper

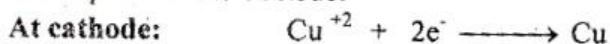
Impure copper is refined by the electrolytic method in the electrolytic cell.

Impure copper acts as anode and a pure copper plate acts as cathode as shown in figure 7.8. Copper sulphate solution in water is used as an electrolyte.

Oxidation reaction takes place at the anode. Copper atoms from the impur copper lose electrons to the anode and dissolve in solution as copper ions:

At anode: $Cu \longrightarrow Cu^{+2}(aq) + 2e^-$

Reduction reaction takes place at the cathode. The copper ions present in the solution are attracted to the cathode. Where they gain electrons from the cathode and become neutral and deposit on the cathode.

At cathode: $Cu^{+2} + 2e^- \longrightarrow Cu$

In the process impure copper is eaten up and purified copper atoms deposit on the cathode.

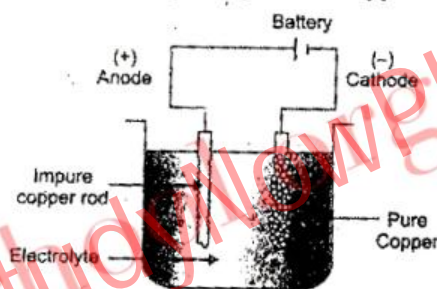



Fig. 7.8 Refining of copper in an electrolytic cell.

Effect of Al_2O_3 and Fe_2O_3 on parent metal

Aluminium has a great tendency to corrosion. However, aluminium corrosion is aluminium oxide (Al_2O_3), a very hard material that actually protects the aluminium from further corrosion. Aluminium oxide corrosion also looks a lot more like aluminium, so it isn't as easy to notice.

When iron corrodes the color changes and it actually expands. This expanding and color change can produce large red flakes that we all know as rust. Unlike aluminium oxide, the expanding and flaking of rust in iron exposes new metal to further rusting.

Q.No.16 Write a note on Chemistry of photography.

Chemistry of photography

In early nineteenth century photographers produced crude images using papers covered with silver nitrate or silver chloride. The exposure of light on photographic plate initiated chemical reaction. The light exposed portion became dark depending the amount or time of exposure. That exposed plate was later on developed to show the image. Those early days "photographs" darkened with time because of ongoing chemical reaction on them. Later on procedures were developed to make the image permanent by use of mercury vapors, followed by washing with sodium hyposulfite ($Na_2S_2O_3$). It dissolved away the silver iodide from the unexposed portion of the plate, and stopped the reaction further. Although technologically more advanced, the basic procedures developed originally are still used in all silver-based photography today.

EXERCISE

MCQ'S

1. **Spontaneous chemical reactions take place in:**
   (a) Electrolytic cell (b) Galvanic cell (c) Nelson's cell, (d) Downs cell
2. **Formation of water from hydrogen and oxygen is:**
   (a) Redox reaction (b) Acid-base reaction
   (c) Neutralization (d) Decomposition
3. **Which one of the following is not an electrolytic cell?**
   (a) Downs cell (b) Galvanic cell (c) Nelson's cell (d) Both a and c
4. **The oxidation number of chromium in $K_2Cr_2O_7$ is:**
   (a) +2 (b) +6 (c) +7 (d) +14
5. **Which one of the following is not an electrolyte?**
   (a) Sugar solution (b) Sulphuric acid solution
   (c) Lime solution (d) Sodium chloride solution
6. **The most common example of corrosion is:**
   (a) Chemical decay (b) Rusting of iron
   (c) Rusting of aluminium (d) Rusting of tin
7. **Nelson's cell is used to prepare caustic soda along with gases. Which of the following gas is produced at cathode:**
   (a) Cl_2 (b) H_2 (c) O_3 (d) O_2
8. **During the formation of water from hydrogen and oxygen, which of the following does not occur:**
   (a) Hydrogen has oxidized (b) Oxygen has reduced
   (c) Oxygen gains electrons (d) Hydrogen behaves as oxidizing agent
9. **The formula of rust is:**
   (a) $Fe_2O_3.nH_2O$ (b) Fe_2O_3 (c) $Fe(OH)_3.nH_2O$ (d) $Fe(OH)_3$
10. **In the redox reaction between Zn and HCl, the oxidizing agent is:**
    (a) Zn (b) H^+ (c) Cl^- (d) H

ANSWR KEY

1	b	2	a	3	d	4	b	5	a	6	b	7	b
8	d	9	a	10	b	K	I	P	S	A	C	A	D E M Y

SHORT QUESTIONS

Q.1 Define oxidation in terms of electrons. Give an example.

Ans: Oxidation is the loss of electron by an atom or an ion e.g,

$$Zn_{(s)} \rightarrow Zn^{+2}_{aq} + 2e^-$$

$$Fe^{+2}_{aq} \rightarrow Fe^{+3}_{aq} + e^-$$

Q.2 Define reduction in terms of 10 s or gain of oxygen or hydrogen. Give an example.

Ans: Reduction is defined as:

"The addition of hydrogen or removal of oxygen during a chemical reaction."

Q.3 What is difference between valency and oxidation state?

Ans:

Valency	Oxidation
• The apparent charge on an atom, ion or molecule which is called valency.	• The apparent charge assigned to an atom of an element in a molecule or an ion.
• Valency is written as the sign followed by the number as valency=OH^{-2}	• No sign

Q.4 Differentiate between oxidizing and reducing agents

Ans:

Oxidizing agents	Reducing Agent
• These are the substance that reduce itself and oxidize other.	• These are the substance that oxidize itself and reduce other.
• They gain the electrons.	• They loss the electrons.
• Non-metals are good oxidizing agent	• Metals are good reducing agents.
• They are more electronegative in nature.	• They are mostly electropositive.

Q.5 Differentiate between strong and weak electrolytes.

Ans:

Strong electrolyte	Weak electrolyte
Definition The electrolyte which ionize completely in solution and produce more ions, are called strong electrolyte.	**Definition** The electrolytes which ionize to a small extent when dissolve in water and could not produce more ions are called weak electrolytes.
Example strong electrolytes are aqueous of NaCl, NaOH, and H_2SO_4 $NaOH_s \xrightarrow{H_2O} Na^+_{aq} + OH^-_{aq}$	Example Weak electrolyte are the aqueous solutions of acetic acid or $Ca(OH_2)$ $CH_3COOH_\ell + H_2O_\ell \rightarrow CH_3COO^-_{aq} + H_3O^+$

Q.6 How electroplating of tin on steel is carried out?

Ans: It involves the dipping of the clean sheet of ion in a both of molten tin and then passing it through pair of rollers.

Q.7 Why steel is plated with nickel before the electroplating of chromium.

Ans: The steel is usually plated first with nickel or copper then by chromium because it does not adhere well on the steel surface. Moreover, it allows moisture to pass through it and metal is-stripped off.

Q.8 How can you explain, that following reaction is oxidation in terms of increase of oxidation number Al" $Al^{\circ} \longrightarrow Al^{+3} + 3e^{-}$

Ans: According to the definition of oxidation:

"The loss of electron is called oxidation" in the above reaction.

"Al° with zero oxidation state changes to Al^{+3} mean it losses 3 electron, so, the oxidation reaction takes place.

Q.9 How can you prove with an example that conversion of an ion to an atom is an oxidation process?

Ans: Example

$$2Cl^{-} \xrightarrow{Oxidation} Cl_2 + 2e^{-}$$

In the above reaction, loss of electron takes place, chloride ion, (anion) convert into chlorine molecule by losing electrons. So oxidation takes place

Q.10 Why does the anode carries negative charge in galvanic cell but positive charge in electrolytic cell? Justify with comments.

Ans: In case of galvanic cell, anode release the electrons, that gathered at anode and create negative charge while in case of electrolytic cell, the anode attached to the positive terminal of the battery, that is why, it carry positive charge.

Q.11 Where do the electrons flow from Zn electrode in Daniel's cell?

Ans: In Daniel cell, the electrons takes flow from Zn electrode (anode) towards the cathode made up of copper.

Q.12 Why do electrodes get their names 'anode' and cathode in galvanic cell?

Ans: In galvanic cell, oxidation takes place at anode while reduction takes place at cathode. And oxidation always takes place at anode while reduction always takes place at cathode.

Fig. 7.3 A Daniel Cell

Q.13 What happens at the cathode in a galvanic cell?

Ans: In galvanic cell, reduction takes place at the cathode as:

$$Cu^{+2}_{aq} + 2e^- \rightarrow Cu_s$$

Q.14 Which solution is used as an electrolyte in Nelson's cell?

Ans: Brine (aqueous solution of NaCl called brine) is used as electrolyte in Nelson's cell.

Q.15 Name the by-products produced in Nelson's cell?

Ans: Hydrogen gas (H_2) and chlorine gas (Cl_2) are the by-product of Nelson's cell as

$$2NaCl_{aq} + 2H_2O_l \rightarrow H_{2(g)} + Cl_{2(g)} + 2NaOH_{aq}$$

Q.16 Why galvanizing is done?

Ans: Galvanizing is done to protect the iron against corrosion.

Q.17 Why an iron grill is painted frequently?

Ans: Iron grill is painted frequently because due to presence of oxygen water in air, it become corrode, so to prevent it from rusting, it is painted.

Q.18 Why O_2 is necessary for rusting?

Ans: Oxygen plays important rule in rusting. Electrons reduce the oxygen molecules in the presence of H^+ ions.

$$O_{2(g)} + 4H^+_{aq} + 4e^- \rightarrow 2H_2O$$

Q.19 In electroplating of chromium, which salt is used as an electrolyte?

Ans: Chromium sulphate with few drops of H_2SO_4 acts as electrolyte.

Q.20 Write the redox reaction taking place during the electroplating of chromium?

Ans: At anode:

$$4OH^-_{aq} \rightarrow 2H_2O_l + 4e^-$$

At cathode:

$$Cr^{+3}_{aq} + 3e^- \rightarrow Cr_{(s)}$$

Overall reaction:

$$Cr_2(SO_4)_{3(s)} \xrightarrow{water} Cr^{+3}_{aq} \rightarrow 3SO^{-2}_{4(aq)}$$

Q.21 In electroplating of silver, from where Ag + come and where they deposit?

Ans: In electroplating of silver Ag^+ ion come form anode while they deposit at cathode.

Q.22 What is the nature of electrode used in electrolyting of chromium?

Ans: The electrolyte is the solution of chromium sulphate $Cr_2(SO_4)$ a few drops of H_2SO_4 are added in the electrolyte to prevent its hydrolyses, while the object to be electroplated acts as cathode and anode is made of antimonies lead.

LONG QUESTIONS

Q.1 Describe the rules for assigning the oxidation state

Ans: See the topic of rule for the assigning the oxidation state.

Q.2 Find out the oxidation numbers of the underlined elements in the following compounds.

(a) $Na_2\underline{S_2}O_4$

(b) $K_2\underline{Cr_2}O_7$

(c) $Ag\underline{N}O_3$

(d) $H\underline{N}O_3$

(e) $K\underline{Mn}O_4$

Ans: See the topic examples of oxidation states.

Q.3 How can a non-spontaneous reaction be carried out in an electrolytic cell. Discuss in detail.

Ans: See the topic electrolytic cell.

Q.4 Discuss the electrolysis of water.

Ans: See the topic electrolysis of water

Q.5 Discuss the construction and working of a cell in which electricity is produced.

Ans: See the topic Galvanic cell.

Q.6 How we can prepare NaOH on commercial scale. Discuss its chemistry along with the diagram.

Ans: See the topic Nelson's cell.

Q.7 Discuss the redox reaction taking place in the rusting of iron in detail.

Ans: See the topic rusting of Iron.

Q.8 Discuss, why galvanizing is considered better than that of tin plating.

Ans: See the topic tin electroplating.

Q.9 What is electroplating? Write down procedure of electroplating.

Ans: See the topic electroplating.

Q.10 What is the principle of electroplating? How electroplating of chromium is carried out?

Ans: See the topic chromium electroplating.

CHAPTER

ELECTROCHEMISTRY

MULTIPLE CHOICE QUESTIONS

1. **Symbol of hydronium ion is:**
   (a) H^+ (b) OH^- (c) H_3O^+ (d) none of these
2. **An example of non electrolyte is:**
   (a) glucose (b) aq NaCl (c) HCl (d) H_2SO_4
3. **An example of weak electrolyte is :**
   (a) HNO_3 (b) HCl (c) H_2SO_4 (d) H_2CO_3
4. **During electrolysis _____ takes place at anode.**
   (a) catenation (b) oxidation (c) reduction (d) addition
5. **Which is not the characteristic of electrolyte?**
   (a) cheap (b) conductor (c) easily oxidized (d) soluble in water
6. **In dry cell ___acts as cathode:**
   (a) Zn cup (b) graphite rod (c) paste (d) steel rod
7. **In Zn-Cu galvanic cell, Zn is dipped in:**
   (a) $ZnSO_4$ (b) $Zn(NO_3)_2$ (c) $CuSO_4$ (d) both a & b
8. **In Zn-Cu galvanic cell, Zn is used as:**
   (a) cathode (b) anode (c) electrode (d) all of above
9. **In $Zn + Cu^{+2} \longrightarrow Zn^{+2} + Cu$, Zn is:**
   (a) oxidized (b) reduced (c) redoxed (d) decomposed
10. **Anions are ________ions.**
    (a) positive (b) neutral (c) negative (d) amphoeteric
11. **Cations are _______ ions.**
    (a) negative (b) positive (c) neutral (d) none of these
12. **Electrolysis of NaCl is done in the cell:**
    (a) electrolytic (b) voltaic (c) down's (d) faradays'
13. **Which one is not produced during electrolysis of aqueous NaCl?**
    (a) Na^+ (b) OH^- (c) H_2O (d) Cl^-
14. **In pure water, out of ____only one molecule ionizes.**
    (a) 6×10^6 (b) 6×10^8 (c) 6×10^{16} (d) 6×10^{12}
15. **Which one is weak electrolyte?**
    (a) citric acid (b) carbonic acid (c) tartaric acid (d) all of these

16. **Metallic, ammonium and hydrogen ions carry _____ charges.**
(a) negative (b) neutral (c) positive (d) partial positive

17. **Electrode through which electrons enter the external circuit:**
(a) anode (b) Cathode (c) electrode (d) none of these

18. **Electrode through which electrons leave the external circuit:**
(a) anode (b) Cathode (c) graphite (d) electrolyte

19. **Which one is conductor?**
(a) naphthalene (b) paraffin wax (c) plastic (d) HCl

20. **Which solution is not a conductor:**
(a) NaCl (b) KI (c) $Co(NH_2)^2$ (d) $CuCl_2$

21. **Rods through which electric current enters or leaves the cell:**
(a) protons (b) electrons (c) electrodes (d) all of these

22. **Spontaneous redox reaction produce current in:**
(a) voltaic cell (b) electrolytic cell (c) galvanic cell (d) none of these

23. **In an oxidation reaction electrons are:**
(a) absorbed (b) lost (c) moved (d) increased

24. **In reduction reaction electrons are:**
(a) lost (b) absorbed (c) kept constant (d) all of these

ANSWER KEY

1	c	6	a	11	b	16	c	21	c
2	a	7	a	12	c	17	a	22	d
3	d	8	b	13	c	18	b	23	b
4	b	9	a	14	b	19	d	24	b
5	a	10	c	15	d	20	c	KIPS	

SHORT QUESTIONS

7.1 OXIDATION AND REDUCTION REACTIONS

Q.1 How can you justify that a reaction between magnesium and oxygen is a redox reaction, while the reaction shows only addition of oxygen (oxidation)

$2\,Mg + O_2 \longrightarrow 2\,MgO$

Ans.

(i) $2Mg \xrightarrow{Oxidation} 2Mg^{+2} + 4e^-$

Magnesium atom Magnesium ion

(ii) $O_2 + 4e^- \xrightarrow{Reduction} 2O^{-2}$

Oxygen molecule Oxide ion

(iii)

In above reaction, loss of electron also occur instead of addition of oxygen while gain of electron is called reduction. So in redox reactions oxidation and reduction occur simultaneously.

Q.2 A reaction between carbon and oxygen involved only addition of oxygen (oxidation), but, it is called a redox reaction. Comment on this.

Ans. As the redox reaction

"A chemical reaction in which oxidation and reduction processes involved is called redox reaction."

In case of carbon reactys with oxygen:

$C \xrightarrow{Oxidation} C^{+4} + 4e^-$

Carbon atom Carbon ion

Oxidation reaction;

$O_2 + 4e^- \xrightarrow{Reduction} 2O^{-2}$

Oxygen molecule Oxide ions

$C^{+4} + 2O^{-2} \longrightarrow CO_2$

Loss of electrons is also called oxidation, while gain of electrons called reduction, so the above reaction is the example of redox reaction.

Q.3 Oxidation and reduction proceed simultaneously. Explain, with an example.

Ans. Oxidation and reduction occurs simultaneously in this reaction.

Example: $FeO + CO \longrightarrow Fe + CO_2$

Q.4 Identify which of the following is oxidation or reduction reaction.

(a) $K \longrightarrow K^+ + 1e^-$

(b) $Br + 1e^- \longrightarrow Br^-$

(c) $Cu \longrightarrow Cu^{+2} + 2e^-$

(d) $I^- \longrightarrow I + 1e^-$

(e) $Fe^{+2} \longrightarrow Fe^{+3} + 1e^-$

Ans. (a) Oxidation (Loss of electron)
(b) Reduction (Gain of electron)
(c) Oxidation (Loss of electron)
(d) Oxidation (Loss of electron)
(e) Oxidation (Loss of electron)

Q.5 An element M reacts with another element X to form MX_2. In terms of loss or gain of electrons, identify the element which is oxidized and which is reduced.

Ans. **Reaction:**

$$M + X_2 \longrightarrow MX_2$$

In this reaction, "M" release two electrons and become "M^{+2}" so oxidation reaction is:

$$M \xrightarrow{Oxidation} M^{+2} + 2e^-$$

While "X" gain two electrons, so reduction reaction is

$$2X + 2e^- \xrightarrow{Reduction} 2X^-$$

Overall reaction:
"M" is oxidized.
X is reduced.

Q.6 How can you justify that the following reaction is not only an oxidation reaction but also a complete redox reaction.

$$\mathbf{FeO + CO \longrightarrow Fe + CO_2}$$

Ans.

Ferrous oxide (FeO) transfer its oxygen to carbon monoxide (CO) to form carbon dioxide (CO_2) and Iron (Fe). In this reaction, oxidation and reduction take place simultaneously as:

$$\mathbf{FeO + CO \longrightarrow Fe + CO_2}$$

Q.7 Explain the term oxidation on the basis of electronic concept with an example

Ans. According to Electronic Concept:
"The loss of electrons by an atom or an ion is called oxidation."
Example:

$$Mg \xrightarrow{Oxidation} Mg^{+2} + 2e^-$$

7.2 OXIDATION STATE AND RULES FOR ASSIGNING OXIDATION STATE

Q.1 Find out the oxidation numbers of the following elements marked in bold in the formulae: $Ba_3(PO_4)_2$, $CaSO_4$, $Cu(NO_3)_2$, $Al_2(SO_4)_3$

Ans. **$Ba_3(PO_4)_2$:**

Barium sulphate

(Oxidation no. of "Ba") 3 + (Oxidation no. of "P")2 + (Oxidation no. of oxygen) 8 = 0

(+2)3 + (Oxidation no. of "P")2 + (-2)8 = 0

+6 + (Oxidation no. of "P")2 - 16 = 0

+6 + (Oxidation no. of "P")2 = + 16

(Oxidation no. of "P" = +16 − 6 = 10

Oxidation no. of "P" $= \frac{+10}{2} = +5$

$CaSO_4$:

(Oxidation no. of "Ca") + (Oxidation no. of "S") + (Oxidation no. of oxygen) 4 = 0
(+2) + (Oxidation no. of "S") + (-2)4 = 0
(+2) + (Oxidation no. of "S") - 8 = 0
+2 + (Oxidation no. of "S") = + 8
Oxidation no. of "S" $= \frac{+8}{2}$
= +4

$Cu(NO_3)_2$:

(Oxidation no. of "Cu") + (Oxidation no. of "N") 2+ (Oxidation no. of oxygen)6 = 0
(+2) + (Oxidation no. of "N")2 + (-2)6 = 0
+2 + (Oxidation no. of "N")2 - 12 = 0
+2 + (Oxidation no. of "N")2 = + 2
(Oxidation no. of "N")2 = +12 – 2
= + 10
Oxidation no. of "N" $= \frac{+10}{2}$
= + 5

$Al_2(SO_4)_3$:

(Oxidation no. of "Al")2 + (Oxidation no. of sulphur)3 + (Oxidation no. of oxygen) 12 = 0
(+3)2 + (Oxidation no. of sulphur)3 + (-2)12 = 0
+6 + (Oxidation no. of "S")3 - 24 = 0
+6 + (Oxidation no. of "S")3 = +24
(Oxidation no. of "S")3 = +24-6
= + 18
$= \frac{+18}{3}$
= + 6

Q.2 In a compound MX_3, find out the oxidation number of M and X.

Ans. Compound = MX_3

If we consider that the oxidation no. of M = +3
[Oxidation no. of M] + [oxidation no. of X] 3 = 0
(+3) + [Oxidation no. of "X"]3 = 0
[Oxidation no. of "X"] 3 = - 3
Oxidation no. of "X" $= \frac{-3}{3}$
Oxidation no. of X = -1

Q.3 Why the oxidation number of oxygen in OF_2 is + 2?

As we know that fluorine is more electronegative than oxygen. So more electronegative atom have high oxidation number also as:

[Oxidation no. of oxygen] + 2 (-1) = 0
(O.N of Oxygen) (-2) = 0
O.N of oxygen = +2

Q.4 In H_2S, SO_2 and H_2SO_4 the sulphur atom has different oxidation number. Find out the oxidation number of sulphur in each compound.

Ans.

(i) H_2S:

[O.N of H) x 2 + (O.N of "S")	= 0
(+1) x 2 + (O.N of "S")	= 0
2 + (O.N of "S")	= 0
O.N of sulphur	= -2

(ii) SO_2:

(O.N of sulphur) + (O.N of oxygen) × 2	= 0
(O.N of sulphur) + (-2) × 2	= 0
O.N of sulphur - 4	= 0
O.N of sulphur	= +4

(iii) H_2SO_4:

(O.N of hydrogen) × 2 + (O.N of sulphur) + (O.N of oxygen) × 4 = 0

(O.N of "H") × 2 + (O.N. of "S") + (O.N of "O") × 4 = 0

(+1) 2 + (O.N of S) + (- 2) 4	= 0
+2 + (O.N of S) + (-2) 4	= 0
O.N of sulphur	= + 8 -2
O.N of sulphur	= +6

Q.5 An element with oxidation state +7 gains electrons to be reduced to oxidation state +2. How many electrons did it accept?

Ans. Consider that element is "X"

As from equation:

$$X^{+7} + 5e^- \longrightarrow X^{+2}$$

This element will gain 5 electron to be reduced.

Q.6 An element "X" has oxidation state "0". What will be its oxidation state when it gains three electrons?

Ans. $X_o + 3e^- \longrightarrow X^{-3}$

Its oxidation state will be (-3)

Q.7 If the oxidation state of an element changes from +5 to – 3, Has it been reduced or oxidized? How many electrons are involved in this process?

Ans. For example the element is "X". As the reaction:

$$X^{+5} + 8e^- \longrightarrow X^{-3}$$

This element gain the electrons, so reduction will be occurred.

It will gain 8 electron 5 are used to neutral its oxidation state while due to more 3 electron it will attain -3 charge.

7.3 OXIDIZING AND REDUCING AGENTS

Q.1 In the following reaction, how can you justify that H_2S is oxidized and SO_2 is reduced.

$$SO_2 + 2H_2S \longrightarrow 2H_2O + 3S$$

Ans. in this reaction, H_2S release the electrons while SO2 absorb the electrons as:

Q.2 The reaction between Mn02 and HCl is a redox reaction written as balance chemical equation.

$$MnO_2 + 4HCl \longrightarrow MnCl_2 + 2H_2O + Cl_2$$

Find out:

(a) **The substance oxidized**

Ans. MnO_2 is oxidized and it lose electron.

(b) **The substance reduce.**

Ans. Chlorine is reduced, it gain electrons.

(c) **The substance reduced**

Ans. Cl_2 is acts oxidizing agent.

(d) **The substance which acts as reducing agent.**

Ans. MnO2 acts as reducing agent.

Q.3 The following reactions are redox reaction. Find out the element which has been reduced and the element which has been oxidized?

(a) $Zn + CuSO_4 \longrightarrow ZnSO_4 + Cu$

(b) $Cu + 2AgNO_3 \longrightarrow Cu(NO_3)_2 + 2Ag$

(c) $H_2S + Cl_2 \longrightarrow 2HCl + S$

Ans. (a) Zn is oxidized and Cu is reduced.

(b) Cu is oxidized and Ag is reduced.

(c) S is oxidized and Cl is reduced.

Q.4 Why the following reaction is not a redox reaction. Explain with reasons?

$NaOH + HCl \longrightarrow NaCl + H_2O$

Ans. In this reaction, as

$$NaOH + HCl \longrightarrow NaCl + H_2O$$

There is no oxidation reduction processes takes place. It is an acid base reaction. Non of atom oxidized or reduced, in this reaction.

7.5 ELECTRO CHEMICAL CELLS

Q.1 Why are the strong electrolytes termed as good conductors?

Ans. Strong electrolytes are good conductor because it completely ionized into its ions and it readily produced more ions as:

$HCl_{(aq)} \longrightarrow H^+_{(aq)} + Cl^-_{(aq)}$ and $NaOH_{(aq)} \longrightarrow Na^+_{(aq)} + OH^-_{(aq)}$

Q.2 Does non-electrolytes forms ions in solution?

Ans. Non-electrolytes does not form ions in solution.

Q.3 What is difference between a strong electrolyte and a weak electrolyte?

Ans.

Strong Electrolyte	Weak Electrolyte
Definition: The electrolytes which ionize completely in solution and produce more ions, are called strong electrolytes.	**Definition:** The electrolytes which ionize to a small extent when dissolve in water and could not produce more ions are called weak electrolytes.
Example: Strong electrolytes are aqueous solutions of NaCl, NaOH and H_2SO_4. $NaOH_{(s)} \xrightarrow{H_2O} Na^+_{(aq)} + OH^-_{(aq)}$	**Example:** Weak electrolytes are the aqueous solution of acetic and or $Ca(OH)_2$ $CH_3COOH_{(l)} + H_2O_{(l)} \longrightarrow CH_3COO^-_{(aq)} + H_3O^+_{(aq)}$

Q.4

Q.5 Identify a strong or weak electrolyte among the following compounds:

$CUSO_4$, H_2CO_3, $Ca(OH)_2$, HCl, $AgNO_3$

Ans. $AgNO_3$ and $CuSO_2$ are strong electrolytes while H_2CO_3 and $Ca(OH)_2$ are the example of weak electrolytes.

Q.6 Which force drives the non-spontaneous reaction to take place?

Ans. Non-spontaneous reactions take place in the presence of an external agent. That external agent are electrons that cause electricity. So electric energy helps the non-spontaneous reactions to proceed.

Q.7 Which type of chemical reaction takes place in electrolytic cell?

Ans. In electrolytic cell, non-spontaneous reaction takes place.

Q.8 What type of reaction takes place at anode in electrolytic cell?

Ans. Oxidation takes place at anode. Anode is a positive charge electrode. The atoms on this electrode release electrons as:

$$2Cl^- \longrightarrow 2Cl + 2e^-$$

Q.9 Why the positively charged electrode is called anode in electrolytic cell?

Ans. The positive charged electrode is called anode because it connected to the (+) terminal of the battery and all electron move towards it as

$$2Cl^- \longrightarrow 2Cl + 2e$$

Q.10 In the electrolysis of water, towards which terminal hydronium ions move?

Ans. In the electrolysis of water, H_3O (hydronium) + or H^+ ions move toward the cathode.

Reaction at cathode: $H^+ + 1e^- \longrightarrow [H]$

Q.11 In the electrolysis of water, where is the oxygen produced?

Ans. In the electrolysis of water oxygen is produced at anode as

Reaction at anode: $4OH^- \longrightarrow 2H_2O_{(aq)} + O_{2(g)} + 4\ e-$

Q.12 Towards which electrode of the electrolytic cell moves the cations and what does they do there?

Ans. Cation ions carry (+) charge, they move towards the cathode, in an electrolytic cell. They gain electron at cathode and oxidized.

Q.13 How the half cells of a galvanic cell are connected? What is function of salt bridge?

Ans. The two half cells of the galvanic cell are connected by salt bridge. Salt bridge is used to maintain the flow of ionic current and the electric neutrality.

7.6 ELECTRO CHEMICAL INDUSTRIES

Q.1 Anode of Downs cell is made of a non-metal, what is its name? What is the function of this anode?

Ans. In Downs cell, there is a large block of graphite, which acts as an anode. The Cl^- ions are oxidized there produce Cl_2 gas at anode.

Q.2 Where does the sodium metal is collected in Downs cell?

Ans. Na+ are reduced at cathode and molten Na-metal floats on the denser molten salt mixture from where it is collected in a side tube.

Q.3 What is the name of the by-product produced in the Downs cell?

Ans. Hydrogen gas and chlorine gas are the by-products produced in the Downs cell.

Q.4 Are anodes of Downs cell and Nelson cell made of same element? If yes, what is its name?

Ans. Yes, in both cells anode is made of graphite.

Q.5 What is the shape of cathode in Nelson's cell? Why is it perforated?

Ans. In Nelson's cell, cathode is made of iron and have U-shaped. It is perforated because extra brine or NaOH can be easily separated out.

Q.6 Which 'ions are discharged at cathode in Nelson's cell and what is produced at cathode?

Ans. The H+ ions are discharged at cathode in Nelson's cell. At cathode, sodium hydroxide (NaOH) and hydrogen (H_2) are produced.

7.7 CORROSION AND ITS PREVENTION

Q.1 What is the difference between corrosion and rusting?

Ans.

Corrosion	Rusting
• Corrosion is the general term used for all the metals	• Corrosion of iron is called rusting
• Corrosion of some metals may be stopped	• Rusting is the continuous process.
• It is redox reaction	• It is also redox reaction.

Q.2 What happens to iron in the rusting process?

Ans. During rusting Fe^{+2} formed that spreads through out the surrounding water and react with O_2 to form the salt $Fe_2O_3.nH_2O$ which is called rust.

Q.3 Rusting completes in how many redox reactions?

Ans. Rusting completes in two redox reactions as:

First Redox Reaction:

$$2Fe_{(s)} + O_{2(aq)} + 4H^{+}_{(aq)} \longrightarrow 2Fe^{+2} + 2H_2O$$

Second Redox Reaction:

$$2Fe^{+2} + \frac{1}{2}O_2 + (2+n)H_2O \longrightarrow Fe_2O_3.nH_2O + 4H^{+}$$

Q.4 Explain the role of O_2 in rusting?

Ans. Oxygen plays important role in rusting. Electrons reduce the oxygen molecules in the presence of H+ ions:

Q.5 State the best method for protection of metal from corrosion.

Ans. The best method for protection of metal from corrosion is the coating of highly resistant metal. Corrosion resistant metals like Zn, Sn and Cr are coated on the surface of iron to protect it from corrosion.

Q.6 What do you mean by galvanizing?

Ans. The process of coating a thin layer of zinc on iron is called galvanizing.

Q.7 What is the advantage of galvanizing?

Ans. A big advantage of galvanizing is that zinc protects the iron against corrosion even after the coating surface is broken.

Q.8 Why tin plated iron is rusted rapidly when tin layer is broken?

Ans. When tin layer is broken the iron is exposed to the air and water, a galvanic cell is established and iron rusts rapidly.

Q.9 Name the metal which is used for galvanizing iron?

Ans. Zinc metal is used for galvanizing iron.

7.8 ELECTROPLATING

Q.1 Define electroplating?

Ans. Electroplating is depositing of one metal over the other by means of electrolysis.

Q.2 How electroplating of zinc is carried out?

Ans. The zinc is deposited on the metal by immersing it in a chemical bath containing electrolyte zinc sulphate. A current is applied which results in zinc being deposited on the target metal i.e., cathode.

Q.3 Which material is used to make cathode electroplating?

Ans. The cathode is made of the objects that is to be electroplated e.g., sheet of iron.

Q.4 Why is the anode made up of a metal to be deposited during electrolysis

Ans. Because, when the current passed, the metal from anode dissolved in the solution and metallic ions migrate to the cathode and discharge or deposit on the object. As a result of this charge, a thin layer of metal deposits on the object.

EXERCISE

MCQ'S

1. Spontaneous chemical reactions take place in:

(a) Electrolytic cell (b) Galvanic cell (c) Nelson's cell, (d) Downs cell

2. Formation of water from hydrogen and oxygen is:

(a) Redox reaction (b) Acid-base reaction

(c) Neutralization (d) Decomposition

3. Which one of the following is not an electrolytic cell?

(a) Downs cell (b) Galvanic cell (c) Nelson's cell (d) Both a and c

4. The oxidation number of chromium in $K_2Cr_2O_7$ is:

(a) +2 (b) +6 (c) +7 (d) +14

5. Which one of the following is not an electrolyte?

(a) Sugar solution (b) Sulphuric acid solution

(c) Lime solution (d) Sodium chloride solution

6. The most common example of corrosion is:

(a) Chemical decay (b) Rusting of iron

(c) Rusting of aluminium (d) Rusting of tin

7. Nelson's cell is used to prepare caustic soda along with gases. Which of the following gas is produced at cathode:

(a) Cl_2 (b) H_2 (c) O_3 (d) O_2

8. During the formation of water from hydrogen and oxygen, which of the following does not occur:

(a) Hydrogen has oxidized (b) Oxygen has reduced

(c) Oxygen gains electrons (d) Hydrogen behaves as oxidizing agent

9. The formula of rust is:

(a) $Fe_2O_3.nH_2O$ (b) Fe_2O_3 (c) $Fe(OH)_3.nH_2O$ (d) $Fe(OH)_3$

10. In the redox reaction between Zn and HCl, the oxidizing agent is:

(a) Zn (b) H^+ (c) Cl^- (d) H

ANSWR KEY

1	b	2	a	3	d	4	b	5	a	6	b	7	b
8	d	9	a	10	b	K	I	P	S	A	C	A	D E M Y

SHORT QUESTIONS

Q.1 Define oxidation in terms of electrons. Give an example.

Ans: Oxidation is the loss of electron by an atom or an ion e.g,

$$Zn_{(s)} \rightarrow Zn^{+2}_{aq} + 2e^-$$

$$Fe^{+2}_{aq} \rightarrow Fe^{+3}_{aq} + e^-$$

Q.2 Define reduction in terms of 10 s or gain of oxygen or hydrogen. Give an example.

Ans: Reduction is defined as:

"The addition of hydrogen or removal of oxygen during a chemical reaction."

Q.3 What is difference between valency and oxidation state?

Ans:

Valency	Oxidation
• The apparent charge on an atom, ion or molecule which is called valency.	• The apparent charge assigned to an atom of an element in a molecule or an ion.
• Valency is written as the sign followed by the number as valency=OH^{2-}	• No sign

Q.4 Differentiate between oxidizing and reducing agents

Ans:

Oxidizing agents	Reducing Agent
• These are the substance that reduce itself and oxidize other.	• These are the substance that oxidize itself and reduce other.
• They gain the electrons.	• They loss the electrons.
• Non-metals are good oxidizing agent	• Metals are good reducing agents.
• They are more electronegative in nature.	• They are mostly electropositive.

Q.5 Differentiate between strong and weak electrolytes.

Ans:

Strong electrolyte	Weak electrolyte
Definition The electrolyte which ionize completely in solution and produce more ions, are called strong electrolyte.	**Definition** The electrolytes which ionize to a small extent when dissolve in water and could not produce more ions are called weak electrolytes.
Example strong electrolytes are aqueous of NaCl, NaOH, and H_2SO_4 $NaOH_s \xrightarrow{H_2O} Na^+_{aq} + OH^-_{aq}$	Example Weak electrolyte are the aqueous solutions of acetic acid or $Ca(OH_2)$ $CH_3COOH_\ell + H_2O_\ell \rightarrow CH_3COO^-_{aq} + H_3O^+$

Q.6 How electroplating of tin on steel is carried out?

Ans: It involves the dipping of the clean sheet of ion in a both of molten tin and then passing it through pair of rollers.

Q.7 Why steel is plated with nickel before the electroplating of chromium.

Ans: The steel is usually plated first with nickel or copper then by chromium because it does not adhere well on the steel surface. Moreover, it allows moisture to pass through it and metal is-stripped off.

Q.8 How can you explain, that following reaction is oxidation in terms of increase of oxidation number Al" $Al^{\circ} \longrightarrow Al^{+3} + 3e^{-}$

Ans: According to the definition of oxidation:

"The loss of electron is called oxidation" in the above reaction.

"Al° with zero oxidation state changes to Al^{+3} mean it losses 3 electron, so, the oxidation reaction takes place.

Q.9 How can you prove with an example that conversion of an ion to an atom is an oxidation process?

Ans: Example

$$2Cl^{-} \xrightarrow{Oxidation} Cl_2 + 2e^{-}$$

In the above reaction, loss of electron takes place, chloride ion, (anion) convert into chlorine molecule by losing electrons. So oxidation takes place

Q.10 Why does the anode carries negative charge in galvanic cell but positive charge in electrolytic cell? Justify with comments.

Ans: In case of galvanic cell, anode release the electrons, that gathered at anode and create negative charge while in case of electrolytic cell, the anode attached to the positive terminal of the battery, that is why, it carry positive charge.

Q.11 Where do the electrons flow from Zn electrode in Daniel's cell?

Ans: In Daniel cell, the electrons takes flow from Zn electrode (anode) towards the cathode made up of copper.

Q.12 Why do electrodes get their names 'anode' and cathode in galvanic cell?

Ans: In galvanic cell, oxidation takes place at anode while reduction takes place at cathode. And oxidation always takes place at anode while reduction always takes place at cathode.

Fig. 7.3 A Daniel Cell

Q.13 What happens at the cathode in a galvanic cell?

Ans: In galvanic cell, reduction takes place at the cathode as:

$$Cu^{+2}_{aq} + 2e^- \rightarrow Cu_s$$

Q.14 Which solution is used as an electrolyte in Nelson's cell?

Ans: Brine (aqueous solution of NaCl called brine) is used as electrolyte in Nelson's cell.

Q.15 Name the by-products produced in Nelson's cell?

Ans: Hydrogen gas (H_2) and chlorine gas (Cl_2) are the by-product of Nelson's cell as

$$2NaCl_{aq} + 2H_2O_l \rightarrow H_{2(g)} + Cl_{2(g)} + 2NaOH_{aq}$$

Q.16 Why galvanizing is done?

Ans: Galvanizing is done to protect the iron against corrosion.

Q.17 Why an iron grill is painted frequently?

Ans: Iron grill is painted frequently because due to presence of oxygen water in air, it become corrode, so to prevent it from rusting, it is painted.

Q.18 Why O_2 is necessary for rusting?

Ans: Oxygen plays important rule in rusting. Electrons reduce the oxygen molecules in the presence of H^+ ions.

$$O_{2(g)} + 4H^+_{aq} + 4_{e^-} \rightarrow +2H_2O$$

Q.19 In electroplating of chromium, which salt is used as an electrolyte?

Ans: Chromium sulphate with few drops of H_2SO_4 acts as electrolyte.

Q.20 Write the redox reaction taking place during the electroplating of chromium?

Ans: At anode:

$$4OH^-_{aq} \rightarrow 2H_2O_l + 4e^-$$

At cathode:

$$Cr^{+3}_{aq} + 3e^- \rightarrow Cr_{(s)}$$

Overall reaction:

$$Cr_2(SO_4)_{3(s)} \xrightarrow{water} Cr^{+3}_{aq} \rightarrow 3SO^{-2}_{4(aq)}$$

Q.21 In electroplating of silver, from where Ag + come and where they deposit?

Ans: In electroplating of silver Ag^+ ion come form anode while they deposit at cathode.

Q.22 What is the nature of electrode used in electrolyting of chromium?

Ans: The electrolyte is the solution of chromium sulphate $Cr_2(SO_4)$ a few drops of H_2SO_4 are added in the electrolyte to prevent its hydrolyses, while the object to be electroplated acts as cathode and anode is made of antimonies lead.

LONG QUESTIONS

Q.1 Describe the rules for assigning the oxidation state

Ans: See the topic of rule for the assigning the oxidation state.

Q.2 Find out the oxidation numbers of the underlined elements in the following compounds.

(a) $Na_2\underline{S_2}O_4$

(b) $K_2\underline{Cr_2}O_7$

(c) $Ag\underline{N}O_3$

(d) $H\underline{N}O_3$

(e) $K\underline{Mn}O_4$

Ans: See the topic examples of oxidation states.

Q.3 How can a non-spontaneous reaction be carried out in an electrolytic cell. Discuss in detail.

Ans: See the topic electrolytic cell.

Q.4 Discuss the electrolysis of water.

Ans: See the topic electrolysis of water

Q.5 Discuss the construction and working of a cell in which electricity is produced.

Ans: See the topic Galvanic cell.

Q.6 How we can prepare NaOH on commercial scale. Discuss its chemistry along with the diagram.

Ans: See the topic Nelson's cell.

Q.7 Discuss the redox reaction taking place in the rusting of iron in detail.

Ans: See the topic rusting of Iron.

Q.8 Discuss, why galvanizing is considered better than that of tin plating.

Ans: See the topic tin electroplating.

Q.9 What is electroplating? Write down procedure of electroplating.

Ans: See the topic electroplating.

Q.10 What is the principle of electroplating? How electroplating of chromium is carried out?

Ans: See the topic chromium electroplating.